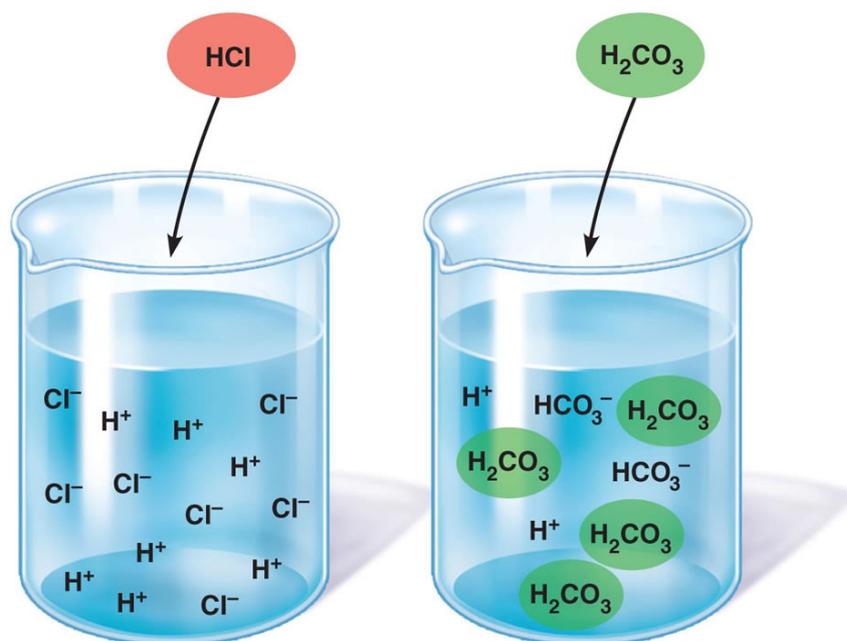# Advanced Higher Chemistry (Revised)

Unit 2 - Physical Chemistry

Weak Acids & Bases

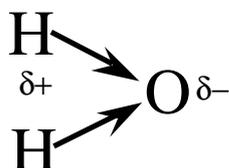

(a) A strong acid such as HCl dissociates completely into its ions.

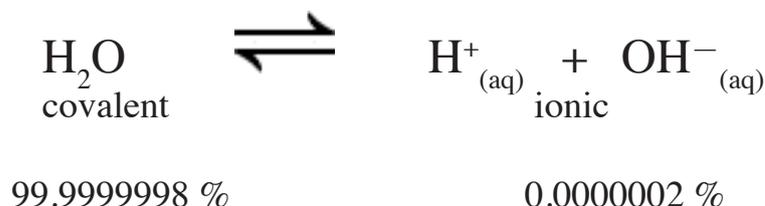(b) A weak acid such as H_2CO_3 does *not* dissociate completely.

Primer Notes (stuff that used to be covered in Higher Course)

Water Equilibrium

This activity considers the equilibrium reactions that take place between the covalent and ionic forms of water.
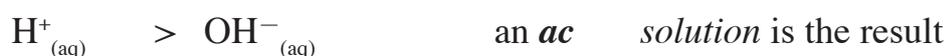
Water molecules have a tendency to 'react' to form $H^+_{(aq)}$ and $OH^-_{(aq)}$ ions. This reaction is, however, ***rev***   so an *equilibrium **mix***   exists.

$$H_2O \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

covalent   ionic

99.9999998 %   0.0000002 %

The ***equi***   *lies well over to the **le***   , though there are still enough ***io***   present to make water a reasonably good ***con***   at *high voltages*.

In *pure water* the ***num***   of $H^+_{(aq)}$ and $OH^-_{(aq)}$ ions are always the ***sa***   , and water is ***neutral***. When other chemicals are added to water, this ***equi***   can be *disturbed* resulting in

$$H^+_{(aq)} > OH^-_{(aq)}$$ an ***ac***   *solution* is the result

$$OH^-_{(aq)} > H^+_{(aq)}$$ an ***alk***   *solution* is the result

Much of this Topic will deal with ***Equilibria in Solutions***.

Strong & Weak Acids

This lesson is about what strong and weak acids are about and why their properties differ.

Definitions

This activity links the terms strong acid and weak acid with what happens when an acid solution is formed.

Hydroch   *acid*, along with ***sulph***   and ***nit***   , remain our '*main 3 acids*'. From now on, they are our '*main 3 **str***   acids*'. So what does the word "*strong*" tell you about an acid? All acids start off as (*polar*) ***cov***   substances.

$$HCl_{(g)} + water \longrightarrow HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$
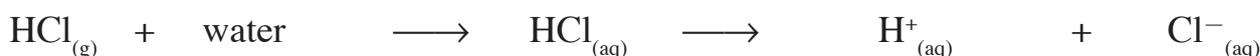
When *dissolved* in water, the ***cov***   acid molecules *dissociate* (split up) into ***io***   . With ***str***   *acids* this *conversion* is 100% and a ***sin***   *arrow* $\longrightarrow$ is correct.

***A strong acid** is one that is **completely dissociated** into ions in solution*

Eth _____ *acid*, along with other ***carb*** _____ *acids*, will be our '*main **we** ____ acids*'. So what does the word *"weak"* tell you about an acid? Again all acids start off as (*polar*) ***cov*** _____ substances.
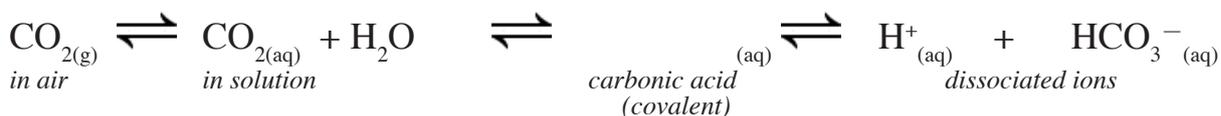
$$CH_3COOH_{(l)} + water \longrightarrow CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$$

ethanoate ions

When *dissolved* in water, very few of the ***cov*** _____ acid molecules *dissociate* (split up) into ***io*** ____ . With ***we*** ____ acids this *conversion* is usually less than 1% . In fact the ***rev*** _____ *reaction* dominates so an ***equi*** _____ *mixture* is formed that contains *very few **io*** ____ compared to ***cov*** _____ *molecules*. *The equilibrium lies well over to the **le*** ____ .

> A ***weak acid*** *is one that is only **partially dissociated** into ions in solution*
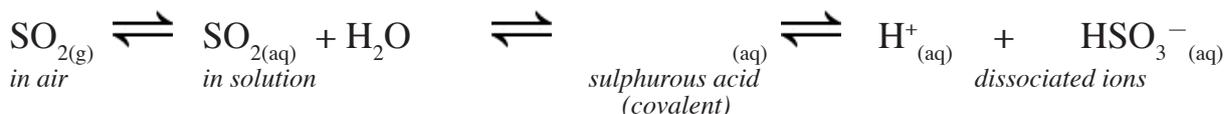
Other examples of ***we*** ____ *acids* are less '*obvious*' and rely more on knowledge gained over the Standard grade and Higher courses. For example,

CO_2 is a ***sol*** ____ gas that (like other ***non-me*** ____ *oxides*) will *dissolve/react* with ***wa*** __ to _____ produce an ***ac*** ____ *solution*.
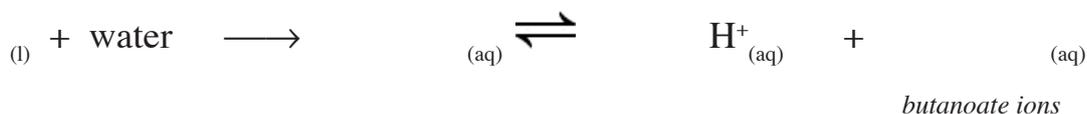
Solutions of CO_2 (think of fizzy drinks) lose CO_2 quickly when open to the air i.e. even _____ the dissolving of CO_2 is a ***rev*** _____ *reaction*.

$$CO_{2(g)} \rightleftharpoons CO_{2(aq)} + H_2O \rightleftharpoons \underset{(aq)}{} \rightleftharpoons H^+_{(aq)} + HCO_3{}^-_{(aq)}$$

in air ___ *in solution* ___ *carbonic acid* ___ *dissociated ions*
(covalent)

Similarly for SO_2 :

$$SO_{2(g)} \rightleftharpoons SO_{2(aq)} + H_2O \rightleftharpoons \underset{(aq)}{} \rightleftharpoons H^+_{(aq)} + HSO_3{}^-_{(aq)}$$

in air ___ *in solution* ___ *sulphurous acid* ___ *dissociated ions*
(covalent)

Most of the time, however, you can expect to deal mainly with ***eth*** _____ , ***prop*** _____ , and ***but*** _____ acid.

$$\underset{(l)}{} + water \longrightarrow \underset{(aq)}{} \rightleftharpoons H^+_{(aq)} + \underset{(aq)}{}$$

propanoate ions

$$\underset{(l)}{} + water \longrightarrow \underset{(aq)}{} \rightleftharpoons H^+_{(aq)} + \underset{(aq)}{}$$

butanoate ions
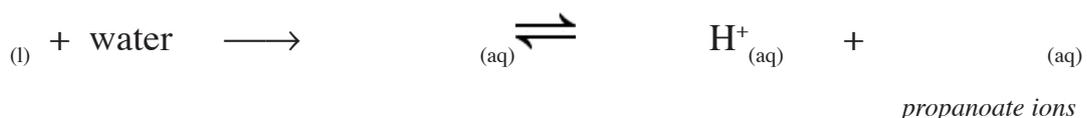
Comparing Strong & Weak Acids	*This activity compares three properties of hydrochloric acid with the same three properties of ethanoic acid*
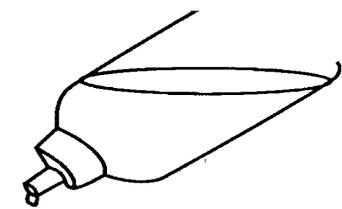
Both acids are ***equi***   - they contain exactly the ***same num***
of *moles* (***same num***   of *molecules*) per litre of solution.

Both acids are ***monoprotic*** (meaning that they are *capable* of
releasing one ***hydr***   ***ion*** per *molecule*).

Comparison	*0.1 M HCl*	*0.1 M CH$_3$COOH*
pH		
Conductivity		
Reaction Rate		

Discussion of Results

pH   there were *equal **num***   of molecules of acid dissolved in both
solutions and 100% of the ***HCl*** molecules ***diss***   to produce
many H$^+$ - ***very l***   ***pH***

Less than 1% of the *CH$_3$COOH* molecules ***diss***   so there were
much fewer H$^+$ ions present - pH still < 7, but ***hi***   than HCl.

Conductivity   there were *equal **num***   of molecules of acid dissolved in both
solutions and 100% of the ***HCl*** molecules ***diss***   to produce
many H$^+$ - ***very hi***   ***conductivity***

Less than 1% of the *CH$_3$COOH* molecules ***diss***   so there were
much fewer H$^+$ ions present - conducts, but *not as good as HCl*.

Reaction Rate   there were *equal **num***   of molecules of acid dissolved in both
solutions and 100% of the ***HCl*** molecules ***diss***   to produce
many H$^+$ - ***very fa***   ***reaction***
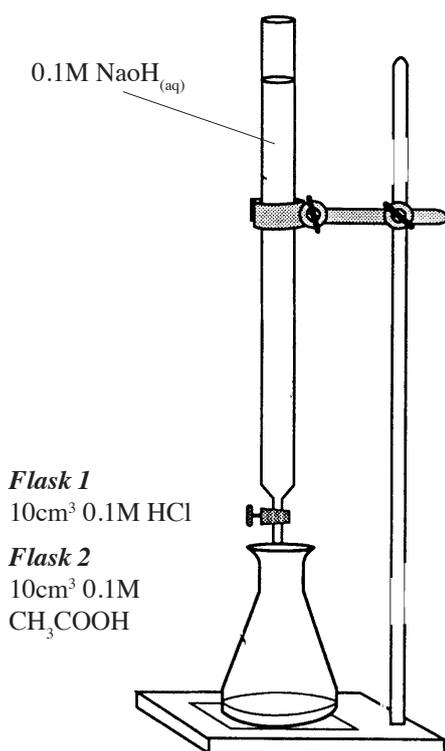
Less than 1% of the *CH$_3$COOH* molecules ***diss***   so there were
much fewer H$^+$ ions present - *fewer **coll***   with magnesium ribbon
so *rate* of ***reaction is slo***   .

Conclusion

*Weak acids produce less H$^+$ ions than strong acids and, therefore, are
less acidic than strong acids **of equal concentration***

| **Titrating Strong &** | *This activity compares the amounts of alkali required to* |
| **Weak Acids** | *neutralise equal quantities of a strong and weal acid* |

0.1M NaoH$_{(aq)}$

Flask 1
10cm^3 0.1M HCl

Flask 2
10cm^3 0.1M
CH$_3$COOH

Vol. needed to **neut**
10 cm^3 0.1M HCl = cm^3

Vol. needed to **neut**
10 cm^3 0.1M CH$_3$COOH = cm^3

Perhaps surprisingly, it takes **exactly the same am** of NaOH to **neut** a **str** acid and a **we** acid.

Both acids were the **same conc** so they both contained the **same num** **of molecules** to begin with.

The strong acid will have **diss** **completely** meaning that **all** the H$^+_{(aq)}$ ions were available to react with the OH$^-_{(aq)}$ ions added from the **bur** .

The weak acid is only **part** **dissociated** so less than 1% of the H$^+_{(aq)}$ ions are available at the beginning.

Weak acid means **rev** reactions; CH$_3$COOH$_{(aq)}$ $\rightleftharpoons$ H$^+_{(aq)}$ + CH$_3$COO$^-_{(aq)}$

ethanoate ions

However, as the H$^+_{(aq)}$ ions react with OH$^-_{(aq)}$ they are *effectively removed* from the **equi** **mixture**. This *slows down* or even *stops* the **rev** **reaction**, but the **for reaction continues**:

$$CH_3COOH_{(aq)} \longrightarrow H^+_{(aq)} + CH_3COO^-_{(aq)}$$

More H$^+_{(aq)}$ ions will be produced which will then react with the OH$^-_{(aq)}$ ions and so on until *every* single CH$_3$COOH$_{(aq)}$ molecule **diss** to form H$^+_{(aq)}$ ions.

Overall, *a weak acid can supply exactly the same number of H$^+_{(aq)}$ as an equal quantity of a strong acid*.

HCl$_{(aq)}$ + NaOH$_{(aq)}$ $\rightarrow$ H$_2$O$_{(l)}$ + NaCl$_{(aq)}$
1 mole *1 mole*

CH$_3$COOH$_{(aq)}$+ NaOH$_{(aq)}$ $\rightarrow$ H$_2$O$_{(l)}$ + NaCH$_3$COO$_{(aq)}$
1 mole *1 mole*

Stoichiometry is the word used to describe the numerical proportions (usually expressed in moles) of substances involved in reactions.

*Strong and weak acids have **different properties**, but they have the **same stoichiometry**.*

Strong and Weak Bases

This lesson is about strong and weak bases and how their properties differ

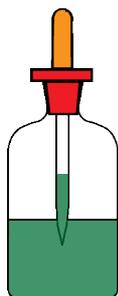**Two Examples of Bases**   *This activity considers whether a base is strong or weak in terms of whether or not it gives an equilibrium mixture in water.*

So *   hydroxide*, along with ***pot*** *   hydroxide* and ***li*** *   hydroxide*, were our '*main 3 alkalis*'. From now on, they are our '*main 3 str   bases*'. So what does the word *"strong"* tell you about a base? Most bases start off as ***io*** *   substances, (oxides or hydroxides)*

$$Na^+OH^-_{(s)} \quad + \quad water \quad \longrightarrow \quad Na^+_{(aq)} \quad + \quad OH^-_{(aq)}$$

When *dissolved* in water, the ***bases*** *dissociate* (split up) into separate ***io*** *   . With* ***str*** *   bases* this *conversion* is 100% and a ***sin*** *   arrow* $\longrightarrow$ is correct.

> *A **strong base** is **completely dissociated** to form OH⁻ ions in solution*

Less ***sol*** *   hydroxides* (*Data Book*) are unlikely to ***diss*** *   comp*   (except at *very low* ***conc*** *   *) and will be considered *weak bases*:

    Group 1   metal hydroxides   -   very soluble
    Group II   metal hydroxides   -   slightly soluble
    Transition metal hydroxides   -   insoluble

The *most important* ***we*** *   ba* *   that you* *must* .... repeat, ***must*** know about is ***amm*** *   *, NH_3. Similar to *weak acids*, ***amm*** *   is a* ***po*** *   cov* *   molecule* that *dissolves readily* in water (***fou*** *   experiment*). Most of the ***amm*** *   * (99.6%) *remains as molecules*, but a small amount *reacts with water* to produce ***hydr*** *   ions. This reaction is* ***rev*** *   *.

$$NH_{(g)} + water \longrightarrow NH_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

   ammonia   *ammonia solution*   *ammonium hydroxide solution*

Bottles can be labelled '*ammonia solution*' or '*ammonium hydroxide solution*' - they are both the ***same mix*** *   . Since the* ***equi*** *   lies well* *over to the le* * , the more correct label would be '*ammonia solution*'

> *A **weak base** is only **partially dissociated** to form OH⁻ ions in solution*

Other examples of ***we*** *   ba* *   would be the* *organic* equivalents of ***amm*** *   - the* *amines*

$$CH_3NH_{2(l)} + water \longrightarrow CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3^+{}_{(aq)} + OH^-_{(aq)}$$

methylamine   *methylamine solution*   *methylammonium hydroxide solution*

Comparing pH & Conductivity

This activity compares the pH and conductivity of equimolar solutions of a strong base and a weak base

Both bases are *equi*   - they contain exactly the *same num*   of *moles* (*same num*   of *molecules/units*) per litre of solution.

Both bases are *monobasic* (meaning that they are *capable* of

Comparison	0.1 M NaOH	0.1 M NH_3
pH		
Conductivity		

Discussion of Results

pH   there were *equal num*   of molecules/units of base dissolved in both solutions and 100% of the *NaOH* units *diss*   to produce *many* OH^- - *very hi*   *pH*

Less than 1% of the NH_3 molecules *diss*   so there were *much fewer* OH^- ions present - pH still >7, but *lo*   than NaOH

Conductivity   there were *equal num*   of molecules/units of base dissolved in both solutions and 100% of the *NaOH* units *diss*   to produce *many* OH^- - *hi*   *conductivity*

Less than 1% of the NH_3 molecules *diss*   so there were *much fewer* OH^- ions present - conducts, but *not as good as NaOH*.

Conclusion

Weak bases produce less OH^- ions than strong bases and, therefore, are less alkaline than strong bases of equal concentration

Stoichiometry of Reactions

This activity compares the amounts of precipitate produced with iron (III) nitrate solution using equal amounts of a strong and a weak base.

The equations for the two reactions will look very similar:

$$Fe^{3+}(NO_3^{-})_{3\,(aq)} \quad + \quad 3\ Na^+OH^-_{(aq)} \quad \rightarrow \quad Fe^{3+}(OH^-)_{3\,(aq)} + 3\ Na^+NO_3^{-}_{(aq)}$$

$$Fe^{3+}(NO_3^{-})_{3\,(aq)} \quad + \quad NH_4^+OH^-_{(aq)} \quad \rightarrow$$

The only difference between these two reactions is that the NaOH has 100% of its OH^- ions available to react, while the ammonia/ammonium hydroxide has less than 0.4%.

10 cm^3 of 1M
$Na^+OH^-_{(aq)}$

10 cm^3 of 1M
$NH_4^+OH^-_{(aq)}$

5 cm^3 of 0.2M $Fe^{3+}(NO_3^-)_3$
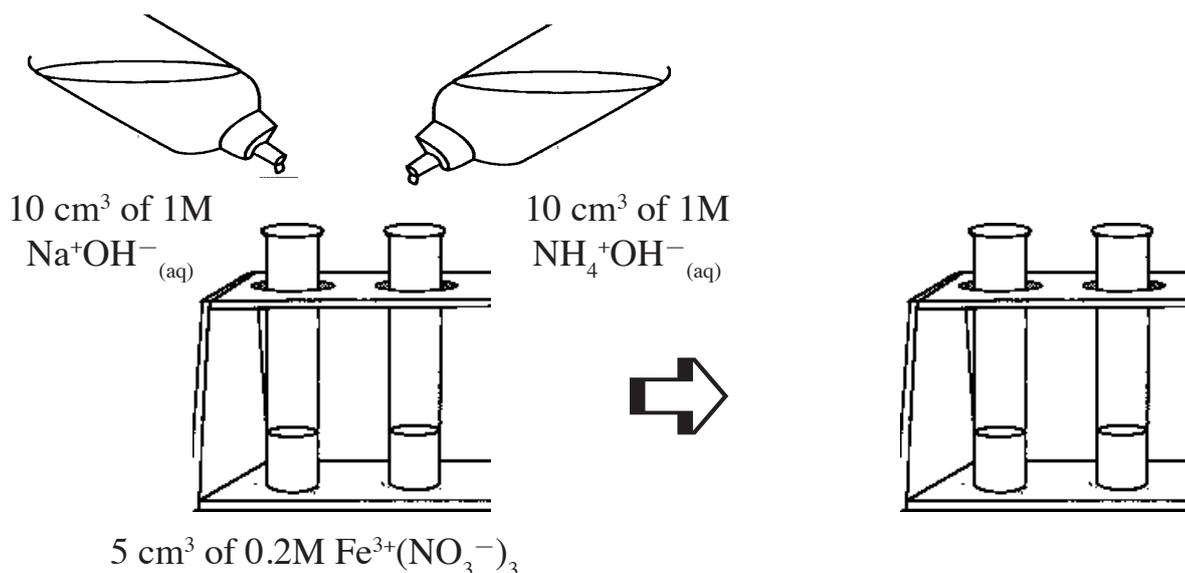
Perhaps surprisingly, **exactly the same am**   of precipitate is produced by a **str**   **base** and a **we**   **base**.

Both bases were the **same conc**   so they both contained the **same num**   **of** **molecules / units** to begin with.

The strong base will have **diss**   **completely** meaning that **all** the $OH^-_{(aq)}$ ions were available to react with the $Fe^{3+}_{(aq)}$ ions .

The weak base is only **part**   **dissociated** so less than   1% of the $OH^-_{(aq)}$ ions are available at the beginning.

However, as the $OH^-_{(aq)}$ ions react with $Fe^{3+}_{(aq)}$ they are *effectively removed* from the **equi**   **mixture**. This *slows down* or even *stops* the **rev**   **reaction**, but the **for** **reaction continues**:

$$NH_{(aq)} \quad + \quad H_2O_{(l)} \longrightarrow \quad NH_4^+_{(aq)} \quad + \quad OH^-_{(aq)}$$

More $OH^-_{(aq)}$ ions will be produced which will then react with the $Fe^{3+}_{(aq)}$ ions and so on until *every* single $NH_{(aq)}$ molecule **diss**   to form $OH^-_{(aq)}$ ions.

Overall, *a weak base can supply exactly the same number of $OH^-_{(aq)}$ as an equal quantity of a strong base*
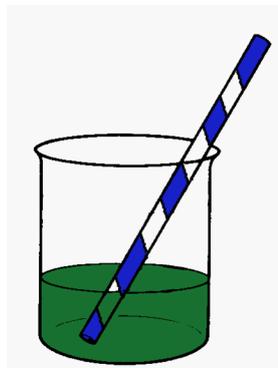
Strong and weak bases have **different properties**, *but they have the* **same stoichiometry**.

Aqueous Solutions

This lesson looks at the effect on pH of dissolving oxides and salts in water

Solutions of Covalent Oxides

This activity examines how it is possible for covalent oxides to affect the pH of water
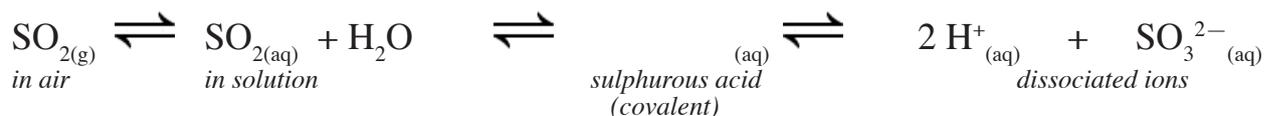
When you **bl** CO_2 into **wa**   the pH will **dr**   from pH =   to about pH =

This is because CO_2 **rea**   with **wa**   to produce the **we**   **acid** called **car**   **acid**. Most of the **car**   **acid** molecules produced remain as **cov**   **molecules** - only a few **diss**   **ions** are produced - so the **position of equi**   is well over to the **le**   .

$$CO_{2(g)} \rightleftharpoons CO_{2(aq)} + H_2O \rightleftharpoons \underset{\substack{carbonic\ acid \\ (covalent)}}{}_{(aq)} \rightleftharpoons 2\,H^+_{(aq)} + CO_3{}^{2-}_{(aq)}$$

in air   *in solution*   *dissociated ions*

Similarly, SO_2 dissolves to produce the weak acid, **sulph**   **acid**, (acid rain).

$$SO_{2(g)} \rightleftharpoons SO_{2(aq)} + H_2O \rightleftharpoons \underset{\substack{sulphurous\ acid \\ (covalent)}}{}_{(aq)} \rightleftharpoons 2\,H^+_{(aq)} + SO_3{}^{2-}_{(aq)}$$

in air   *in solution*   *dissociated ions*

Other **cov**   oxides are either **inso**   (CO, NO) or dissolve to form **str**   **acids** ($NO_2 \rightarrow nitric\ acid$), ($SO_3 \rightarrow sulphuric\ acid$), ($P_2O_5 \rightarrow phosphoric\ acid$)
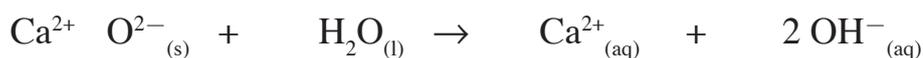
Solutions of Ionic Oxides

This activity examines how soluble ionic oxides are able to affect the pH of water
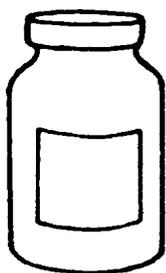
When you add CaO to **wa**   the pH will **ri**   from pH =   to about pH =

This is because CaO also **re**   with **wa**   to from *calcium* **hydr**   .

$$Ca^{2+}\ O^{2-}_{(s)} + H_2O_{(l)} \rightarrow Ca^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

This reaction is *not rev*   , so *calcium* **hydr**   is considered a **str**   **base**. However, $Ca(OH)_2$ is *not very* **sol**   so if too much is produced, some of it will **prec**   out as **so**   :

$$Ca(OH)_{2\ (aq)} \rightleftharpoons Ca(OH)_{2\ (s)}$$

Only the *Group 1 oxides/hydroxides* can be relied upon to be **sol**   at *high* **conc**   . *Group II oxides/hydroxides* are only **sol**   at *low* **conc**   . *Transition metal oxides/ hydroxides* are **insol**   (DATA BOOK)

Solutions of Salts

This activity looks at the effect that dissolving salts can have on the pH of water

Salts are ***io***   ***compounds*** derived from an ***ac***   and a ***ba***   . Though we *describe* the reaction between an ***ac***   and a ***ba***   as ***neutr***   , it would be wrong to *assume* that the final ***sa***   *solution* is always ***neu***   .

It depends on the ***stre***   of the *parent ac*   and the *parent ba*   .

Salt	Formula	Parent Base & Strength	Parent Acid & Strength	pH and type of solution
sodium chloride	Na^+Cl^-	sodium hydroxide *STRONG*	hydrochloric acid *STRONG*	pH = 7 *NEUTRAL*
sodium sulphite		sodium hydroxide	sulphurous acid	pH =
ammonium nitrate		ammonia	nitric acid	pH =
magnesium sulphate		magnesium hydroxide	sulphuric acid	pH =
sodium stearate	$Na^+C_{17}H_{35}COO^-$	sodium hydroxide	stearic acid	pH =
ammonium sulphate		ammonia	sulphuric acid	pH =

From the results above, 3 rules can be made:

❶   ***ACIDIC*** solutions   when parent ***ACID STRONG***, parent *BASE WEAK*

❷   ***ALKALI*** solutions   when parent *ACID WEAK*, parent ***BASE STRONG***

❸   *NEUTRAL* solutions   when parent ***ACID STRONG***, parent ***BASE STRONG***

These rules can be used to ***PREDICT*** the pH of a salt solution.

To ***EXPLAIN*** the pH of a salt solution is more complicated and these explanations come next.

Equilibria in Salt Solutions

This topic examines how it is possible for salts to have an effect on the pH of water

Sodium Ethanoate Solution

This activity explains why salts of strong bases and weak acids form alkaline solutions

*Sodium ethanoate is the salt made from **sod   hyd   (str   base**) and **eth   acid** (**we   acid**) and you would **predict** that it would produce an **alk   solution***

$$\textbf{\textit{STRONG \qquad WEAK}}$$
$$\textbf{\textit{BASE \qquad ACID}}$$

$$\textbf{\textit{predict}} \quad pH \quad 7$$

But how do we *explain* this?

*Sodium ethanoate is **sol   , and **diss   to release **io   into the **wa   ;***

$$CH_3COONa_{(s)} \quad + \quad water \quad \rightarrow \quad _{(aq)} \quad + \quad Na^+{}_{(aq)}$$

There are, however, already two *rea* taking place in the water;

$$H_2O \quad \rightleftharpoons \quad _{(aq)} \quad + \quad _{(aq)}$$

Secondary reactions between the salt ions and the water ions are possible;

$$CH_3COOH_{(s)} + \quad water \quad \rightarrow \quad \boxed{\;_{(aq)}\;} \quad + \quad Na^+{}_{(aq)}$$

$$H_2O \quad \rightleftharpoons \quad _{(aq)} \quad + \quad OH^-{}_{(aq)}$$

$$\updownarrow$$
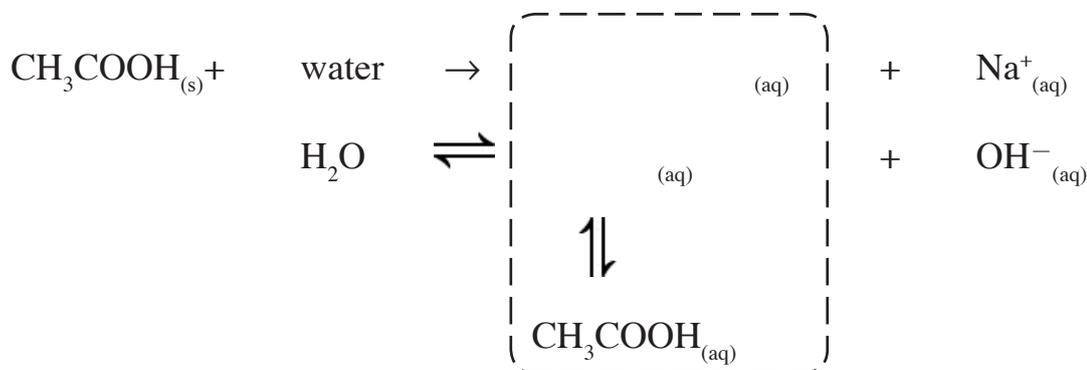
$$CH_3COOH_{(aq)}$$

Eth   ions ($CH_3COO^-{}_{(aq)}$) will react with **hyd   ions** ($H^+{}_{(aq)}$) to produce **cov** *molecules* of **eth   acid** ($CH_3COOH_{(aq)}$). Being a **we   acid**, only a **sm   proportion** of these molecules will **diss   to reform $H^+{}_{(aq)}$ ions.

As a result $H^+{}_{(aq)}$ ion **conc   will **dr   . This will **sl   down** the **back** reaction so for a while;

$$H_2O \quad \rightleftharpoons \quad H^+{}_{(aq)} + \quad OH^-{}_{(aq)}$$

The *extra* $H^+{}_{(aq)}$ ions produced will be *mopped up* by the **eth   ions** ($CH_3COO^-{}_{(aq)}$) but the **extra** $OH^-{}_{(aq)}$ ions build up so,

$$OH^- \quad > \quad H^+ \quad \text{and pH} \quad 7, \text{ alkaline solution}$$

| **Ammonium Chloride Solution** | *This activity explains why salts of weak bases and strong acids form acidic solutions* |

Ammonium chloride is the salt made from ***amm   hyd   (we   base*) and ***hydro***   acid (***str   acid***) and you would ***predict*** an ***ac***   solution
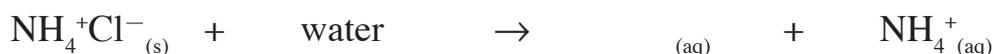
$$\begin{array}{ll} \textit{\textbf{WEAK}} & \textit{\textbf{STRONG}} \\ \textit{\textbf{BASE}} & \textit{\textbf{ACID}} \end{array} \qquad \textit{\textbf{predict}} \quad \text{pH} \quad 7$$

But how do we *explain* this?

Ammonium chloride is ***sol***   , and ***diss***   to release *io*   into the *wa*   ;

$$NH_4^+Cl^-_{(s)} \quad + \quad \text{water} \quad \rightarrow \quad _{(aq)} \quad + \quad NH_4^+{}_{(aq)}$$

There are, however, already two ***rea***   taking place in the water;

$$H_2O \quad \rightleftharpoons \quad _{(aq)} \quad + \quad OH^-{}_{(aq)}$$

Secondary reactions between the salt ions and the water ions are possible;

$$NH_4^+Cl^-_{(s)} \quad + \quad \text{water} \quad \rightarrow \quad _{(aq)} \quad + \quad \boxed{\begin{array}{c} NH_4^+{}_{(aq)} \\ \\ OH^-{}_{(aq)} \\ \updownarrow \\ NH_{3(aq)} \; + \; H_2O_{(l)} \end{array}}$$

$$H_2O \quad \rightleftharpoons \quad _{(aq)} \quad +$$
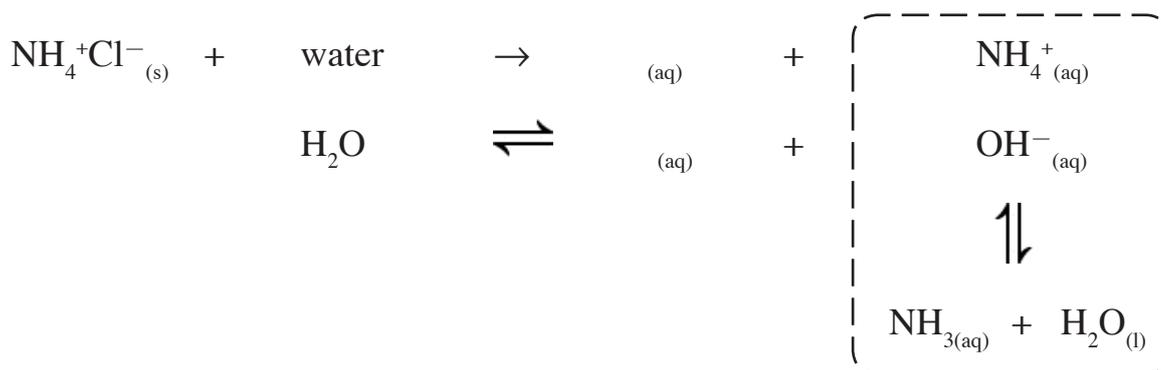
Amm   *ions* ($NH_4^+{}_{(aq)}$) will react with ***hydr***   *ions* ($OH^-{}_{(aq)}$) to produce ***cov*** *molecules* of ***amm***   ($NH_{3(aq)}$) and ***water*** ($H_2O_{(l)}$). Being a ***we***   *base*, only a ***sm*** *proportion* of these molecules will ***diss***   to reform $OH^-{}_{(aq)}$ ions.

As a result $OH^-{}_{(aq)}$ ion ***conc***   will ***dr***   . This will ***sl***   ***down*** the ***back*** reaction so for a while;

$$H_2O \quad \rightleftharpoons \quad H^+{}_{(aq)} + \quad OH^-{}_{(aq)}$$

The *extra* $OH^-{}_{(aq)}$ ions produced will be *mopped up* by the ***amm***   *ions* ($NH_4^+{}_{(aq)}$) but the ***extra*** $H^+{}_{(aq)}$ ions build up so,

$$H^+ \; > \; OH^- \qquad \text{and pH} \quad 7, \; \text{acidic solution}$$

| **Sodium Chloride Solution** | *This activity considers the changes taking place when the salt of a strong base and a strong acid is added to water* |

Sodium chloride is the salt made from **sod   hyd   (str   base)** and **hydro** acid (**str   acid**) and you would **predict** that it would produce a **neu   ** solution

$$\begin{array}{ccc} \textbf{\textit{STRONG}} & \textbf{\textit{STRONG}} & \\ \textbf{\textit{BASE}} & \textbf{\textit{ACID}} & \textbf{\textit{predict}} \quad \text{pH} \quad 7 \end{array}$$

But how do we *explain* this?

Sodium chloride is **sol   **, and **diss   ** to release **io   ** into the **wa   ** ;

$$Na^+Cl^-_{(s)} \quad + \quad water \quad \rightarrow \quad _{(aq)} \quad + \quad Cl^-_{(aq)}$$
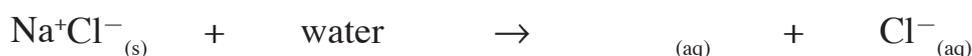
Again there are already two **rea   ** taking place in the water;

$$H_2O \quad \rightleftharpoons \quad _{(aq)} \quad + \quad OH^-_{(aq)}$$

Secondary reactions, however, between the salt ions and the water ions are **not** possible;

$$Na^+Cl^-_{(s)} \quad + \quad water \quad \rightarrow \quad \boxed{\begin{array}{c} _{(aq)} \\ _{(aq)} \\ \uparrow \end{array}} \quad + \quad \boxed{\begin{array}{c} Na^+_{(aq)} \\ OH^-_{(aq)} \\ \uparrow \\ Na^+OH^- \end{array}}$$

$$H_2O \rightleftharpoons$$
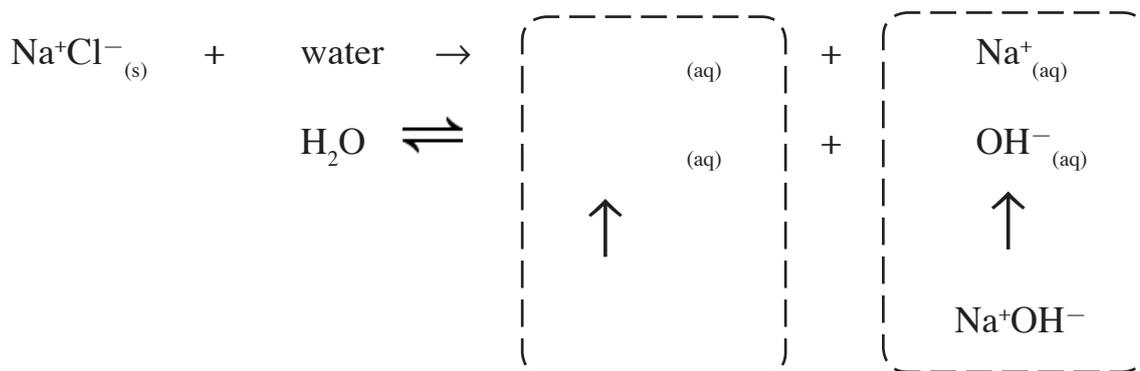
If $H^+_{(aq)}$ ions were to react with $Cl^-_{(aq)}$ they would form molecules of $HCl_{(aq)}$ but , being a strong acid, these molecules would dissociate 100% so all the $H^+_{(aq)}$ ions would return.

If $OH^-_{(aq)}$ ions were to react with $Na^+_{(aq)}$ they would form NaOH but , being a strong base, would dissociate 100% so all the $OH^-_{(aq)}$ ions would return.

$$H^+ \;=\; OH^- \qquad and \; pH \qquad 7, \; neutral \; solution$$

> *When asked to **explain** the pH of salt solutions, you will be expected to produce the two equations involved and show which ions react.*

Concentration & pH

This lesson considers the relationship between the concentration of ions in solution and the pH value

Hydrogen Ions & pH
This activity investigates the numerical relationship between the hydrogen ion concentration and pH value

pH is a ***mea***   of the ***conc***   of $H^+_{(aq)}$ ions in a ***sol***   .

The ' *p*' in pH stands for '*power of*' and '*H*' is, of course, '*Hydrogen*'

Square brackets, [ ] , are used in Chemistry for '*concentration of*', so
[$H^+_{(aq)}$] = *concentration of* $H^+_{(aq)}$ *ions*

Concentration of acid *(in terms of molarity)*	*Concentration of $H^+_{(aq)}$* *(as a power of 10)*	*pH* *(to nearest whole number)*
0.1 M	10^{-1} mol ℓ^{-1}	1
0.01 M	10 mol ℓ^{-1}	
0.001 M	10 mol ℓ^{-1}	
0.0001 M	10 mol ℓ^{-1}	
0.00001 M	10 mol ℓ^{-1}	

From the table, it can be seen that the power of 10 index corresponds to the pH value

$$[H^+] = 10^{-\text{index}} \qquad pH = \text{index}$$

For example,

an acid of pH = 3   means $[H^+]$ = 10 mol ℓ^{-1}

water of pH = 7   means $[H^+]$ = 10 mol ℓ^{-1}

alkali of pH = 12   means $[H^+]$ = 10 mol ℓ^{-1}

Notice that, even when we deal with alkalis, we continue to measure $[H^+]$.

acids $[H^+]$ > in water   ie, $[H^+]$ > 10 mol ℓ^{-1}

alkalis $[H^+]$ < in water   ie, $[H^+]$ < 10 mol ℓ^{-1}

As you've seen, there is a simple relationship between [H⁺] and pH;

$$0.1 \text{ mol } \ell^{-1} \text{ HCl} \quad = \quad 10^{-1} \text{ mol } \ell^{-1} [H^+] \quad = \quad \text{pH of } 1$$
$$1.0 \text{ mol } \ell^{-1} \text{ HCl} \quad = \quad 10 \quad \text{mol } \ell^{-1} [H^+] \quad = \quad \text{pH of}$$
$$10 \text{ mol } \ell^{-1} \text{ HCl} \quad = \quad 10 \quad \text{mol } \ell^{-1} [H^+] \quad = \quad \text{pH of}$$
$$100 \text{ mol } \ell^{-1} \text{ HCl} \quad = \quad 10 \quad \text{mol } \ell^{-1} [H^+] \quad = \quad \text{pH of}$$

In reality, there are no acids that are *sol*   enough to allow 100 moles of acid to dissolve in 1 litre of water. ***Conc***   HCl is about 12 mol ℓ^{-1} while ***conc*** H_2SO_4 is about 20 mol ℓ^{-1} .

Similarly 1 mol ℓ^{-1} and 10 mol ℓ^{-1} ***alk***   solutions are possible but *sol*   limits prevent ***alk***   solutions of higher ***conc***   .

Therefore, our *pH scale* is really;

$$... \text{-}1 \ \ 0 \ \ 1 \ \ 2 \ \ 3 \ \ 4 \ \ 5 \ \ 6 \ \ 7 \ \ 8 \ \ 9 \ \ 10 \ \ 11 \ \ 12 \ \ 13 \ \ 14 \ \ 15 \ ...$$

Ionic Product for Water

This activity considers the relationship between the concentrations of hydrogen ions and hydroxide ions in solutions of different pH values.

The water ***equi***   is a ***dyn***   *equilibrium*, both reactions taking place all the time; in ***pu***   water, ***ac***   solutions and ***alk***   solutions.

$$H_2O \ \rightleftharpoons \ H^+_{(aq)} + \ OH^-_{(aq)}$$

If an ***ac***   is ***diss***   in water, the [H⁺] will be ***inc***   . This will *speed up* the ***back***   reaction. Some of the added H⁺ ions will be *converted* into H_2O but the [H⁺] will still be ***hi***   than in ***pu***   water. However, OH⁻ ions are *converted* into H_2O.   ***So [H⁺]↑ but [OH⁻]↓***

$$H_2O \ \rightleftharpoons \ H^+_{(aq)} + \ OH^-_{(aq)}$$

If a ***ba***   is ***diss***   in water, the [OH⁻] will be ***inc***   . This will *speed up* the ***back***   reaction. Some of the added OH⁻ ions will be *converted* into H_2O but the [OH⁻] will still be ***hi***   than in ***pu***   water. However, H⁺ ions are *converted* into H_2O.   ***So [OH⁻]↑ but [H⁺]↓***

$$H_2O \ \rightleftharpoons \ H^+_{(aq)} + \ OH^-_{(aq)}$$

Clearly, there is always going to be a *relationship between [H⁺] and [OH⁻].*

In ***pu***   water the pH is   , so   [H⁺] = 10   mol ℓ^{-1} and   [OH⁻] = 10   mol ℓ^{-1}

Multiplying the two ***conc***   gives what is called the ***Ionic Product*** for water

ionic product = [H⁺] x [OH⁻] = 10   mol ℓ^{-1} x 10   mol ℓ^{-1} =   (units?)

[NaOH]	pH measured	[H⁺] calculated	[OH⁻]	[H⁺] x [OH⁻] ionic product
0.0001 M	10	10^{-10} mol l^{-1}	10 mol l^{-1}	10 mol² l^{-2}
0.01 M	12	10 mol l^{-1}	10 mol l^{-1}	10 mol² l^{-2}
1.0 M	14	10 mol l^{-1}	10 mol l^{-1}	10 mol² l^{-2}

From the table above, it can be seen that the ionic product is a constant; its value remains unchanged in a solution of any pH.

$$\textit{ionic product} = [H^+] \; x \; [OH^-] = 10 \quad \text{mol}^2 \, l^{-2}$$

Calculations Involving the Ionic Product

This activity uses the ionic product of water to calculate the concentrations of hydrogen or hydroxide ions in solutions

Example 1. *In a solution of pH 9 what is the concentration of $OH^-_{(aq)}$?*

Example 2. *If $[OH^-_{(aq)}]$ is 0.01 mol l^{-1} what is   a) the concentration of hydrogen ions*
                                                          and   b) the pH value?

Example 3. *If a very concentrated solution of NaOH contains 40g in 100 cm³ of solution,*

calculate   a) $[OH^-_{(aq)}]$

                        b) $[H^+_{(aq)}]$

and   c)   *the pH*

Example 4. *Very concentrated hydrochloric acid can contain 1 mol in 100 cm³ solution.*

        a)   *What is the pH value?*

        b)   *Calculate the hydroxide ion concentration.*

Acids & Bases

Acids and Bases

6. *The **pH scale** is a continuous range from **below** 0 to **above** 14.*

0.1 M HCl = pH 1,   1M = pH 0,   10M = pH -1
0.1 M NaOH = pH 13,   1M = pH 14,   10M = pH 15

7. ***Integral** (whole number) values from pH 0 to 14 can be related to concentrations of $H^+_{(aq)}$ in mol / l.*

0.1 M $H^+_{(aq)}$ $= 10^{-1}$ $=$ pH 1
0.01 M $H^+_{(aq)}$ $= 10^{-2}$ $=$ pH 2
0.001 M $H^+_{(aq)}$ $= 10^{-3}$ $=$ pH 3

8. *In water and aqueous solutions with a pH value of 7, the concentrations of $H^+_{(aq)}$ and $OH^-_{(aq)}$ are both 10^{-7} mol / l at 25°C.*

0.0000001 M $H^+_{(aq)}$ $= 10^{-7}$ $=$ pH 7
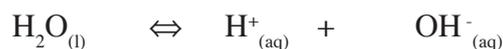
9. *The concentration of $H^+_{(aq)}$ and $OH^-_{(aq)}$ can be calculated from the concentration of the other by using $[H^+_{(aq)}] [OH^-_{(aq)}] = 10^{-14}$ mol / l.*

$[H^+_{(aq)}] = $ ' concentration of ' $H^+_{(aq)}$

pH = 12;   so $[H^+_{(aq)}] = 10^{-12}$

$\quad 10^{-12}$ x $[OH^-_{(aq)}] = 10^{-14}$ so $[OH^-_{(aq)}] = 10^{-2}$

10. *In water and aqueous solutions there is an equilibrium between $H^+_{(aq)}$ and $OH^-_{(aq)}$ and water molecules.*

$H_2O_{(l)}$ $\Leftrightarrow$ $H^+_{(aq)}$ + $OH^-_{(aq)}$

Strong and Weak

11. *In aqueous solutions, **strong** acids are **completely dissociated** but **weak** acids are only **partially dissociated.***

HCl $\rightarrow$ $H^+_{(aq)}$ + $Cl^-_{(aq)}$ (100%)
$H_2CO_3 \Leftrightarrow$ $H^+_{(aq)}$ + $HCO_{3\,(aq)}^-$ (< 1%)
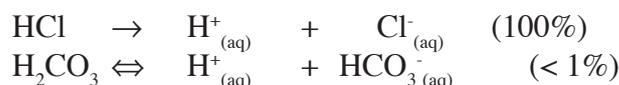
12. *Equimolar solutions of weak and strong acids differ in pH, conductivity and reaction rates but not in stoichiometry (amount) of reactions.*

pH -
$\quad$ 0.1 M HCl,   pH = 1   (10^{-1} $H^+_{(aq)}$)
$\quad$ 0.1 M CH_3COOH,   pH = 4   (10^{-4} $H^+_{(aq)}$)
conductivity -
$\quad$ 0.1 M HCl,   100mA   (10^{-1} $H^+_{(aq)}$)
$\quad$ 0.1 M CH_3COOH,   15mA   (10^{-4} $H^+_{(aq)}$)
rates -
$\quad$ 0.1 M HCl + Mg,   fast   (10^{-1} $H^+_{(aq)}$)
$\quad$ 0.1 M CH_3COOH + Mg,   slow   (10^{-4} $H^+_{(aq)}$)
stoichiometry (amounts) -
$\quad$ 25cm³ 0.1 M HCl   + 25cm³ 0.1 M NaOH
$\quad$ 25cm³ 0.1 M CH_3COOH + 25cm³ 0.1 M NaOH

13. *The weakly acidic nature of **ethanoic acid**, **sulphur dioxide** and **carbon dioxide** can be explained by reference to equations showing the equilibrium.*
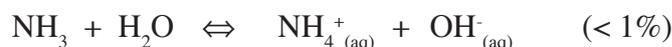
$CH_3COOH_{(l)} \rightarrow CH_3COOH_{(aq)} \Leftrightarrow H^+_{(aq)} + CH_3COO^-_{(aq)}$
$CO_{2(aq)} + H_2O_{(l)} \rightarrow H_2CO_3 \Leftrightarrow H^+_{(aq)} + HCO^-_{3\,(aq)}$
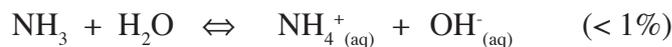$SO_{2(aq)} + H_2O_{(l)} \rightarrow H_2SO_3 \Leftrightarrow H^+_{(aq)} + HSO^-_{3\,(aq)}$
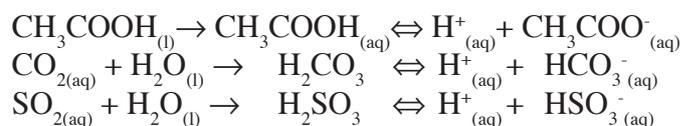
14. *In aqueous solutions, strong bases are **completely dissociated** but **weak** bases are only **partially dissociated.***

$NaOH$ $\rightarrow$ $Na^+_{(aq)}$ + $OH^-_{(aq)}$ (100%)
$NH_3 + H_2O$ $\Leftrightarrow$ $NH_4^+_{(aq)}$ + $OH^-_{(aq)}$ (< 1%)
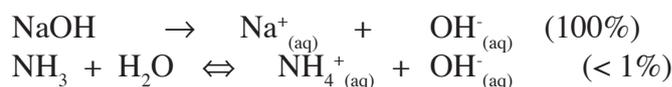
15. *Equimolar solutions of weak and strong bases differ in pH, conductivity but not in stoichiometry (amount) of reactions.*

16. *The weakly alkaline nature of **ammonia**, can be explained by reference to an equation showing the equilibrium.*

$$NH_3 + H_2O \Leftrightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)} \qquad (< 1\%)$$

$$NH_3 + H_2O \Leftrightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)} \qquad (< 1\%)$$

pH of salt solutions

17 *A soluble salt of a **strong acid** and a **strong base** dissolves in water to produce a **neutral** solution.*

18. *A soluble salt of a **weak acid** and a **strong base** dissolves in water to produce an **alkaline** solution*

19. *A soluble salt of a **strong acid** and a **weak base** dissolves in water to produce an **acidic** solution*

Salt	Acid	Base	pH
eg NaCl	strong	strong	= 7
NaCH$_3$COO	weak	strong	> 7
NH$_4$Cl	strong	weak	< 7
NH$_4$CH$_3$COO	weak	weak	?

20. ***Soaps** are salts of weak acids and strong bases.*

21. *The acidity, alkalinity or neutrality of the above kinds of salt can be **explained** by reference to the appropriate **equilibria** (plural !).*

$$H_2O_{(l)} \Leftrightarrow \boxed{H^+{}_{(aq)}} + OH^-{}_{(aq)}$$
$$Na_2SO_3 \rightarrow \boxed{SO_3^-{}_{(aq)}} + Na^+{}_{(aq)}$$

salt

$$\boxed{\Updownarrow}$$

$$\boxed{HSO_3^-{}_{(aq)}}$$ a weak acid so will not

$$\boxed{(or\ H_2SO_3)}$$ fully ionise; **H$^+$ ions removed**